

## Key to all reactions involved in ion tests

### Cations:

Ion	Tests	Observations
$\text{Ag}^+$	Add 2 drops $\text{NaOH}_{(\text{aq})}$ Add 2 drops $\text{NH}_3_{(\text{aq})}$ to a new sample Add excess $\text{NH}_3_{(\text{aq})}$	<ul style="list-style-type: none"> <li>forms a brown precipitate</li> <li>forms a brown precipitate</li> <li>precipitate dissolves</li> </ul>
Name and formula of precipitate: <b>silver hydroxide</b> Equation for formation of precipitate: $\text{Ag}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})} \rightarrow \text{AgOH}_{(\text{s})}$ Equation for dissolving: $\text{AgOH}_{(\text{s})} + 2\text{NH}_3_{(\text{aq})} \rightarrow [\text{Ag}(\text{NH}_3)_2]\text{OH}_{(\text{aq})}$ or $[\text{Ag}(\text{NH}_3)_2]^+ + \text{OH}^-$		
$\text{Al}^{3+}$	Add 2 drops $\text{NaOH}_{(\text{aq})}$ Add excess $\text{NaOH}_{(\text{aq})}$ Add 2 drops $\text{NH}_3_{(\text{aq})}$ to a new sample Add excess $\text{NH}_3_{(\text{aq})}$ Add $\text{H}_2\text{SO}_4_{(\text{aq})}$ to a new sample	<ul style="list-style-type: none"> <li>forms a white precipitate</li> <li>precipitate dissolves</li> <li>forms a white precipitate</li> <li>precipitate remains</li> <li>forms a colourless solution</li> </ul>
Name and formula of precipitate: <b>aluminium hydroxide, <math>\text{Al}(\text{OH})_3</math></b> Equation for formation of precipitate: $\text{Al}^{3+}_{(\text{aq})} + 3\text{OH}^-_{(\text{aq})} \rightarrow \text{Al}(\text{OH})_3_{(\text{s})}$ Equation for dissolving: $\text{Al}(\text{OH})_3_{(\text{s})} + \text{OH}^-_{(\text{aq})} \rightarrow [\text{Al}(\text{OH})_4]^-_{(\text{aq})}$ Reason for adding sulfuric acid: <b>to check that it isn't lead (lead sulfate is insoluble). A more reliable test would be to add potassium iodide; a bright yellow precipitate indicates that the cation is lead and no precipitate would indicate zinc or aluminium (assuming it was amphoteric)</b>		

<b>Ba<sup>2+</sup></b>	Add 2 drops NaOH <sub>(aq)</sub> Add excess NaOH <sub>(aq)</sub> Add H <sub>2</sub> SO <sub>4(aq)</sub> to a new sample	<ul style="list-style-type: none"> <li>• forms a white precipitate</li> <li>• precipitate remains</li> <li>• forms a white precipitate</li> </ul>
Name and formula of first precipitate: <b>barium hydroxide, Ba(OH)<sub>2</sub></b> Equation for formation of precipitate: <b>Ba<sup>2+</sup><sub>(aq)</sub> + 2OH<sup>-</sup><sub>(aq)</sub> → Ba(OH)<sub>2</sub>(s)</b> Name and formula of second precipitate: <b>barium sulfate; BaSO<sub>4</sub></b> Equation for formation of precipitate: $\text{Ba}^{2+}_{(aq)} + \text{SO}_4^{2-}_{(aq)} \rightarrow \text{BaSO}_4(s)$ What other cations could produce similar results to barium, and how are they distinguished: <ul style="list-style-type: none"> <li>- <b>lead (sulfate is insoluble but can be distinguished by amphotericity)</b></li> <li>- <b>calcium and strontium would be difficult to distinguish except by flame test, but you won't be given these (presumably because QA think flame tests are too exciting)</b></li> </ul>		
<b>Cu<sup>2+</sup></b>	Add 2 drops NaOH <sub>(aq)</sub> Add 2 drops NH <sub>3(aq)</sub> to a new sample Add excess NH <sub>3(aq)</sub>	<ul style="list-style-type: none"> <li>• forms a blue precipitate</li> <li>• forms a blue precipitate</li> <li>• forms a deep blue solution</li> </ul>
Name and formula of precipitate: <b>copper hydroxide, Cu(OH)<sub>2</sub></b> Equation for formation of precipitate: <b>Cu<sup>2+</sup><sub>(aq)</sub> + 2OH<sup>-</sup><sub>(aq)</sub> → Cu(OH)<sub>2</sub>(s)</b> Name of deep blue compound and equation for formation: <b>copper II tetrammine complex</b> $\text{Cu(OH)}_2(s) + 4\text{NH}_3(aq) \rightarrow [\text{Cu(NH}_3)_4](\text{OH})_2(aq) \text{ or } [\text{Cu(NH}_3)_4]^{2+} + 2\text{OH}^-$		
<b>Fe<sup>2+</sup></b>	Add 2 drops NaOH <sub>(aq)</sub>	<ul style="list-style-type: none"> <li>• forms a green precipitate</li> </ul>
Name and formula of precipitate: <b>iron II hydroxide (or ferrous hydroxide) Fe(OH)<sub>2</sub></b> Equation for formation of precipitate: <b>Fe<sup>2+</sup><sub>(aq)</sub> + 2OH<sup>-</sup><sub>(aq)</sub> → Fe(OH)<sub>2</sub>(s)</b> If this were a 'forensic' science unit I would make you confirm that it was indeed iron II by reacting it with hydrogen peroxide (to oxidise it to iron III) and then adding thiocyanate to get the red complex. There are other green precipitates (just not on your list) Note that iron II solutions often contain small amounts of iron III because of reaction with atmospheric oxygen. For this reason they will often give a weak colour change with thiocyanate ions.		
<b>Fe<sup>3+</sup></b>	Add 2 drops NaOH <sub>(aq)</sub> Add 2 drops KSCN <sub>(aq)</sub> to a new sample	<ul style="list-style-type: none"> <li>• forms an orange precipitate</li> <li>• forms a dark red solution</li> </ul>
Name and formula of precipitate: <b>iron III hydroxide (or ferric hydroxide) Fe(OH)<sub>3</sub></b> Equation for formation of precipitate: <b>Fe<sup>3+</sup><sub>(aq)</sub> + 3OH<sup>-</sup><sub>(aq)</sub> → Fe(OH)<sub>3</sub>(s)</b> Equation for formation of red solution: $\text{Fe}^{3+}_{(aq)} + \text{SCN}^{-}_{(aq)} \rightarrow [\text{FeSCN}]^{2+}_{(aq)}$		

<b>Mg<sup>2+</sup></b>	Add 2 drops NaOH <sub>(aq)</sub> Add excess NaOH <sub>(aq)</sub> Add H <sub>2</sub> SO <sub>4(aq)</sub> to a new sample	<ul style="list-style-type: none"> <li>• forms a white precipitate</li> <li>• precipitate remains</li> <li>• forms a colourless solution</li> </ul>
Name and formula of precipitate: <b>magnesium hydroxide, Mg(OH)<sub>2</sub></b> Equation for formation of precipitate: <b>Mg<sup>2+</sup><sub>(aq)</sub> + 2OH<sup>-</sup><sub>(aq)</sub> → Mg(OH)<sub>2</sub>(s)</b> Distinguishing features of this cation <ul style="list-style-type: none"> <li>• <b>it ISN'T amphoteric, distinguished from aluminium</b></li> <li>• <b>the sulfate is soluble, distinguished from barium</b></li> </ul>		
<b>Pb<sup>2+</sup></b>	Add 2 drops NaOH <sub>(aq)</sub> Add excess NaOH <sub>(aq)</sub> Add 2 drops NH <sub>3(aq)</sub> to a new sample Add excess NH <sub>3(aq)</sub> Add H <sub>2</sub> SO <sub>4(aq)</sub> to a new sample	<ul style="list-style-type: none"> <li>• forms a white precipitate</li> <li>• precipitate dissolves</li> <li>• forms a white precipitate</li> <li>• precipitate remains</li> <li>• forms a white precipitate</li> </ul>
Name and formula for first precipitate: <b>lead hydroxide, Pb(OH)<sub>2</sub></b> Equation for formation of first precipitate: <b>Pb<sup>2+</sup><sub>(aq)</sub> + 2OH<sup>-</sup><sub>(aq)</sub> → Pb(OH)<sub>2</sub>(s)</b> Equation for dissolving of first precipitate: $\text{Pb(OH)}_2(\text{s}) + 2\text{OH}^-(\text{aq}) \rightarrow [\text{Pb(OH)}_4]^{2-}(\text{aq})$ Name and formula of second precipitate: <b>lead hydroxide, Pb(OH)<sub>2</sub> (again;</b> Equation for formation of second precipitate: <b>Pb<sup>2+</sup><sub>(aq)</sub> + 2OH<sup>-</sup><sub>(aq)</sub> → Pb(OH)<sub>2</sub>(s)</b> What distinguishes lead ions from aluminium or barium? <ul style="list-style-type: none"> <li>• <b>to distinguish from barium - is amphoteric</b></li> <li>• <b>to distinguish from aluminium - sulfate is insoluble</b></li> </ul>		
<b>Na<sup>+</sup></b>	Add 2 drops NaOH <sub>(aq)</sub>	<ul style="list-style-type: none"> <li>• does not form a precipitate</li> </ul>

**Sodium is the 'what it is if it isn't anything else' choice here.**

**Be aware that ammonium is similar to sodium, but has the distinctive ammonium smell (e.g. if you add NaOH)**

**If QA allowed flame tests, we could distinguish sodium and potassium (which are not easily distinguishable by any chemical test)**

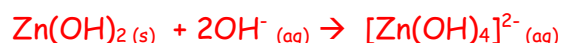
**Remember that solutions labelled either NH<sub>3</sub> or NH<sub>4</sub>OH contain NH<sub>3</sub> , NH<sub>4</sub><sup>+</sup> and OH<sup>-</sup> all present in the one solution. In dilute solutions the two ionic species will predominate, hence the initial precipitation of hydroxide solids on addition of ammonia solution. Once the relative concentration of ammonia passes equivalence the concentration of molecular NH<sub>3</sub> increases rapidly, and thus its availability as a complexing agent. This is the reason that the precipitates redissolve, as the ammino complexes (if they form) are more soluble than the 'naked' cation. However, this can only happen once the concentration of molecular ammonia is sufficiently high.**

<b>Zn<sup>2+</sup></b>	Add 2 drops NaOH <sub>(aq)</sub> Add excess NaOH <sub>(aq)</sub> Add 2 drops NH <sub>3(aq)</sub> to a new sample Add excess NH <sub>3(aq)</sub>	<ul style="list-style-type: none"> <li>• forms a white precipitate</li> <li>• precipitate dissolves in excess</li> <li>• forms a white precipitate</li> <li>• precipitate dissolves in excess</li> </ul>
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Name and formula for first precipitate: **zinc hydroxide, Zn(OH)<sub>2</sub>**

Equation for formation of first precipitate: **Zn<sup>2+</sup><sub>(aq)</sub> + 2OH<sup>-</sup><sub>(aq)</sub> → Zn(OH)<sub>2</sub>(s)**

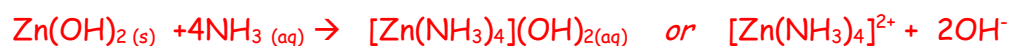
Equation for dissolving of first precipitate:



Name and formula for second precipitate: **zinc hydroxide, Zn(OH)<sub>2</sub>**

Equation for formation of second precipitate: **Zn<sup>2+</sup><sub>(aq)</sub> + 2OH<sup>-</sup><sub>(aq)</sub> → Zn(OH)<sub>2</sub>(s)**

Equation for dissolving of second precipitate:



- **zinc is amphoteric like lead and aluminium, but neither of those form a complex with ammonia**

Ion	Tests	Observations
$\text{Cl}^-$	Add red litmus Add $\text{AgNO}_{3(\text{aq})}$ Add $\text{NH}_{3(\text{aq})}$	<ul style="list-style-type: none"> <li>does not change the colour</li> <li>forms a white precipitate</li> <li>precipitate dissolves</li> </ul>
Name and formula of precipitate: <b>silver chloride, <math>\text{AgCl}</math></b> Equation for formation of precipitate: $\text{Ag}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})} \rightarrow \text{AgCl}_{(\text{s})}$ Equation for precipitate dissolving $\text{AgCl}_{(\text{s})} + 4\text{NH}_{3(\text{aq})} \rightarrow [\text{Ag}(\text{NH}_3)_4]^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$		
$\text{CO}_3^{2-}$	Add red litmus Add $\text{HCl}_{(\text{aq})}$	<ul style="list-style-type: none"> <li>turns red litmus blue</li> <li>bubbles are produced</li> </ul>
Name of gas in bubbles: <b>carbon dioxide, <math>\text{CO}_2</math></b> Equation for formation of gas bubbles: $2\text{H}^+_{(\text{aq})} + \text{CO}_3^{2-}_{(\text{aq})} \rightarrow \text{H}_2\text{O}_{(\text{l})} + \text{CO}_{2(\text{g})}$		
$\text{I}^-$	Add red litmus Add $\text{AgNO}_{3(\text{aq})}$ Add $\text{NH}_{3(\text{aq})}$	<ul style="list-style-type: none"> <li>does not change the colour</li> <li>forms a yellow precipitate</li> <li>precipitate remains</li> </ul>
Name and formula of precipitate: <b>silver iodide, <math>\text{AgI}</math></b> Equation for formation of precipitate: $\text{Ag}^+_{(\text{aq})} + \text{I}^-_{(\text{aq})} \rightarrow \text{AgI}_{(\text{s})}$ <b>Note that insolubility of iodide in excess ammonia can be unreliable, as can the yellow colour of the precipitate</b> Specify another test you could do to show the presence of iodide ions; give any relevant equations. <b>Add lead nitrate and you will get a bright yellow precipitate of lead iodide if the anion present is indeed iodide; if it is chloride or another halide there will be no precipitate or it will be a different colour.</b> $\text{Pb}^+_{(\text{aq})} + \text{I}^-_{(\text{aq})} \rightarrow \text{I}_{(\text{s})}$		
$\text{NO}_3^-$	Add red litmus Add $\text{AgNO}_{3(\text{aq})}$ Add $\text{BaCl}_{2(\text{aq})}$	<ul style="list-style-type: none"> <li>does not change the colour</li> <li>does not form a precipitate</li> <li>does not form a precipitate</li> </ul>
$\text{OH}^-$	Add red litmus Add $\text{HCl}_{(\text{aq})}$	<ul style="list-style-type: none"> <li>turns red litmus blue</li> <li>no bubbles are produced</li> </ul>
<b>Note: to confirm for sure it is hydroxide and to have an equation to write (to potentially gain another emergency mark should you need it) an optional test for this would be to add lead nitrate, which you are given as a reagent. Initially, there will be no precipitate because of the excess of hydroxide ions over lead (forming plumbate complex). As the relative concentration of lead ions increases, a precipitate of lead hydroxide would appear. This precipitate could be made to disappear again by adding a little more of the "unknown"</b> <b>If you had an alkaline solution and added lead and got an immediate precipitate, it could be that you had carbonate ions but had missed the bubbles with acid. If you were absolutely sure no bubbles you may have something strange like phosphate or citrate (not on the list).</b>		

$SO_4^{2-}$	Add red litmus Add $AgNO_{3(aq)}$ Add $BaCl_{2(aq)}$	<ul style="list-style-type: none"> <li>• does not change the colour</li> <li>• does not form a precipitate</li> <li>• forms a white precipitate</li> </ul>
Name and formula of precipitate: <b>barium sulfate, <math>BaSO_4</math></b> Equation for formation of precipitate: $Ba^{2+}_{(aq)} + SO_4^{2-}_{(aq)} \rightarrow BaSO_{4(s)}$		

Explain why carbonate and hydroxide turn red litmus blue but the other anions do not:

they are alkaline i.e. have pH of above 7. Sodium carbonate is formed by the reaction of a strong base (sodium hydroxide) with a weak acid (carbonic acid) so has a pH greater than 7. This is why you are not given bicarbonate, as it would not have a pH of above 7 (being further reaction of sodium carbonate with carbonic acid). Other basic salts include the alkali salts of citric, acetic, phosphoric, benzoic and similar acids.

Note that nitrate is the "what it is if it isn't anything else" anion choice, because all nitrates are soluble. This also applies to acetate (ethanoate) but that is not on your list. There is a positive test for nitrate (brown ring test) but it is not required in this paper.